

## Thomson Model of the Atom

J. J. Thomson performed experiments with cathode rays in an attempt to understand electricity – which was still a mystery in the late 1800s.

Thomson’s 1897 Experiments - state the conclusions Thomson drew from each of his famous cathode ray experiments:

1. **First Experiment**: Thomson directed the beam at an electrometer and tried to separate the evidence of charge from the path of the beam. *What connection did Thomson find between charge and the cathode rays? Was the charge positive or negative?*
2. **Second Experiment**: Thomson tried passing the cathode ray through an electric field. *How did cathode ray beam behave when it passed through an electric field? What did he conclude after his second experiment?*
3. **Third Experiment**: Thomson did some careful measurements on how much the path of the cathode ray was bent in a magnetic field and how much energy they carried. From this work Thomson could describe the mass/charge ratio of the cathode ray particles. *What amazing result did Thomson find?*
4. **Thomson’s Atomic Model**: Thomson presented three hypotheses from his experiments. Only two were accepted by physicists – in fact the third was shown to be wrong! From the first two came a model of the atom known as the *Plum Pudding* model. Complete the atom drawing below to illustrate Thomson’s plum pudding model. Explain how this fits with his observations.

## Thomson Model and Sticky Tape

Let’s see how we can use Thomson’s model to explain the behavior of the sticky tape when we made our tape stacks. A few atoms from the *top tape* and the *bottom tape* are represented in the diagram below. Electrons have been added to the atoms before adhesion. Now add electrons to each atom “in contact” and “separated” to show what happens to the electrons when we make a tape stack out of neutral pieces of tape and then pull them apart.

Top after separation (+)

Bottom after separation (-)

Before adhesion In contact Separated

Describe the *macroscopic* changes in the tapes (what you observed when the tape was separated) and then provide a *microscopic* explanation based on Thomson’s model of the atom and your drawings (why it happened).

**Behavior of Foil and Paper with Charged Tapes**

We observed that neither foil (metal atoms) nor paper (non-metal atoms) would attract each other. But foil and paper are **both** attracted to **both** the charged tapes (top and bottom). ***How can we use the pudding model of atoms to explain the differences we observed?***

* Several atoms from the paper and foil are drawn on the next page. The ones on the left have no charged object near them. The ones on the right are next to a top tape (+ charge).
* Add force vectors to the non-metal (paper) atoms and the top tape in the first row to show the attraction between the paper and the tape. Then do the same for the foil and the tape in the second row. Be sure the *size of the vectors* shows the relative strengths of the attractions.
* Now draw the electrons in each atom “bowl” to show their arrangements when no charged object is near present and then when a charged object is brought near.

Top Tape

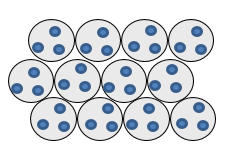
Metal (Al foil)

*no charged object near*

Metal (Al foil)

*top tape near*

Top Tape



Explain why these arrangements of electrons would produce the observed attractions.

**Ionic Compounds**

We have observed evidence that aqueous ionic compounds conduct electricity but the solid form does not. Dalton determined that compounds form from individual atoms of elements in very specific whole number ratios. Thompson determined that those same atoms consist of a positive sphere with negative electrons embedded throughout. We will now examine the patterns that exist for the ratios in which these elements combine in order and why their solid forms do not conduct electricity.

1. Write the formula and draw the particle diagram for each compound.

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| Atoms  involved | 1 lithium  1 fluorine | 1 lithium  1 chlorine | 2 lithium  1 oxygen | 2 lithium  1 sulfur | 3 lithium  1 nitrogen |
| formula |  |  |  |  |  |
| particle diagram |  |  |  |  |  |

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| Atoms  involved | 1 sodium  1 fluorine | 1 sodium  1 chlorine | 2 sodium  1 oxygen | 2 sodium  1 sulfur | 3 sodium  1 nitrogen |
| formula |  |  |  |  |  |
| particle diagram |  |  |  |  |  |

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| Atoms  involved | 1 magnesium  2 fluorine | 1 magnesium  2 chlorine | 1 magnesium  1 oxygen | 1 magnesium  1 sulfur | 3 magnesium  2 nitrogen |
| formula |  |  |  |  |  |
| particle diagram |  |  |  |  |  |

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| Atoms  involved | 1 calcium  2 fluorine | 1 calcium  2 chlorine | 1 calcium  1 oxygen | 1 calcium  1 sulfur | 3 calcium  2 nitrogen |
| formula |  |  |  |  |  |
| particle diagram |  |  |  |  |  |

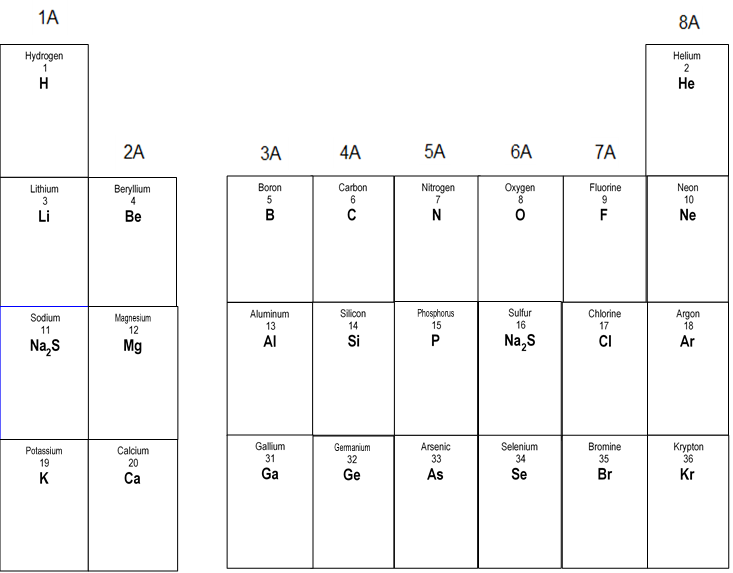
|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| Atoms  involved | 1 gallium  3 fluorine | 1 gallium  3 chlorine | 2 gallium  3 oxygen | 2 gallium  3 sulfur | 3 gallium  2 nitrogen |
| formula |  |  |  |  |  |
| particle diagram |  |  |  |  |  |

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| Atoms  involved | 1 gallium  3 fluorine | 1 gallium  3 chlorine | 2 gallium  3 oxygen | 2 gallium  3 sulfur | 3 gallium  2 nitrogen |
| formula |  |  |  |  |  |
| particle diagram |  |  |  |  |  |

1. Write each formula from Question 1 in the boxes corresponding to its elements. For example, the compound formed from sodium and sulfur have been written in the box for sodium and in the box for sulfur. Now add the rest.

Group 18

Group 1



Group 17

Group 16

Group 15

Group 14

Group 13

Group 2

1. What patterns do you find in the formulas of the compounds formed in the table in #2?
2. What patterns do you find in the formulas of the compounds formed in the table in #2?
3. Based on these patterns, predict the formulas of the compounds formed by the ions below.

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| Atoms involved | **Ratio of ions in compound** | | | | | |
| \_\_\_ potassium  \_\_\_ bromine | \_\_\_ potassium  \_\_\_ selenium | \_\_\_ potassium  \_\_\_ phosphorus | \_\_\_ beryllium  \_\_\_ bromine | \_\_\_ beryllium  \_\_\_ selenium | \_\_\_ beryllium  \_\_\_ phosphorous |
| formula |  |  |  |  |  |  |
| Particle diagram |  |  |  |  |  |  |

1. How does a neutral atom become a positive ion? A negative ion?
2. Recall metals conduct electricity in all phases. Is this because the electrons were mobile or stationary?
3. When the top (+) tape was held near the metal aluminum foil, it attracted the foil by moving the electrons in the foil.
   1. Based on your model, if the top tape were permitted to touch the foil, would the foil gain or lose electrons?
   2. Would the aluminum become positively charged or negatively charged?
   3. If metals attract nonmetals, what charges do nonmetals form?
4. In ionic compounds, the metal particles tend to form positively charged ions (cations), while non-metal particles tend to form negatively charged ions (anions). However, when these same metal and non-metal particles are combined to form compounds they do not conduct electricity as solids. Use your evidence to explain why ions may conduct in the liquid phase and not in the solid phase. Draw a model to help your explanation.
5. Are all cations equal but oppositely charged from all anions? Use the ratio or metals to nonmetals observed in question 1 to help you answer this question.
6. Make the generalizations you can about the charge of the ions formed by elements in groups 1, 2, 13, 15, 16, and 17 based on the ratio of atoms in each of the compounds they form. It might help to look at your particle diagrams in #1 and consider what charges the elements might have in order to result in neutral compounds. Write your generalizations on the periodic table provided.
7. Complete the models below to show how neutral atoms transfer **electrons** to form neutral compounds. Include the charges of the ions in model of the compound.

**Na(s) Cl2(g)🡪 NaCl(s)**

**Ca(s) Cl2(g) 🡪 CaCl2(s)**

**Oxidation Numbers**

Assign oxidation numbers to each element in the following:

1. NaCl Na\_\_\_ Cl\_\_\_
2. Fr2S Fr\_\_\_ S\_\_\_
3. Rb2O Rb\_\_\_ O \_\_\_
4. CaO Ca \_\_\_ O\_\_\_
5. KI K\_\_\_ I\_\_\_
6. AlP Al\_\_\_ P\_\_\_
7. H2 H\_\_\_
8. MgCl2 Mg\_\_\_ Cl\_\_\_
9. I2 I\_\_\_\_
10. BeO Be\_\_\_\_ O\_\_\_\_
11. Na3P Na\_\_\_ P\_\_\_
12. SrF2 Sr\_\_\_\_ F \_\_\_\_
13. AlCl3 Al\_\_\_\_ Cl\_\_\_\_
14. CsF Cs\_\_\_\_ F\_\_\_\_
15. KCl K \_\_\_\_ Cl\_\_\_\_
16. K2O K \_\_\_\_ O\_\_\_\_
17. O3 O \_\_\_\_
18. LiI Li \_\_\_\_ I\_\_\_\_
19. NaBr Na \_\_\_\_ Br\_\_\_\_
20. BaBr2 Ba\_\_\_ Br\_\_\_
21. Al2O3 Al\_\_\_ O\_\_\_
22. RaTe Ra\_\_\_ Te\_\_\_
23. Ga2Se3 Ga\_\_\_ Se\_\_\_
24. Ca3N2 Ca\_\_\_ N\_\_\_

**REDOX Reactions**

A redox reaction is a reaction in which electrons are transferred from one element to another. The reaction involves at least two elements, one that will give up an electron, and one that will receive that electron. The term redox comes from two words, “oxidation” and “reduction.” If something is oxidized, it “burns” in oxygen, as shown below:

Mg(s) + O(g) 🡪 Mg+2 + O-2 🡪 MgO(s)

As you can see, Mg is being oxidized and it loses its two electrons, while oxygen gains them. **Oxidation is defined as the loss of electrons**.

Considering the name of this reaction, if something is being oxidized, the other element must be reducing. Neither reduction nor oxidation can happen alone. That means that oxygen, in this example, is being reduced because it gains the electrons. **Reduction is defined as the gain of electrons**.

A simple way to remember this is to remember that **LEO the lion goes GER**.

**LEO** = loss of electrons is oxidation.

**GER** = gain of electrons is reduction.

Also, Mg and O are “free” elements, which means their oxidation numbers are both 0. Now you can see that **if and element’s oxidation number increases, that element is being oxidized. If an element’s oxidation number decreases it is being reduced (the number is reducing).**

EXAMPLES: Indicate which element is being oxidized, and which is being reduced.

1. Cr3+ + Fe2+  🡪 Cr2+  + Fe3+
2. F2 + O2-  🡪 F1-  + O2
3. Sn + N5+  🡪 Sn4+ + N4+
4. NaCl 🡪 Na+ + Cl-
5. Li2O 🡪 Li + O2
6. Cl2 + KBr 🡪 KCl + Br2

In the following examples state if they are redox reactions and provide evidence.

1. 2Ca + O2  🡪 2CaO
2. Cl2 + 2NaBr 🡪 NaCl + Br2
3. 2LiCl + CaO 🡪 CaCl2 + Li2O

**Redox Half Reactions**

A half reaction shows either the oxidation or reduction portion of a redox equation including if the electrons are gained or lost. A reduction half reaction shows an atom or ion gaining one of more e-: Fe3+ + 3e- 🡪 Fe Notice that the e- is on the left.

An oxidation half reaction shows an atom or ion losing one or more e-:

Mg 🡪 Mg2+ + 2 e- Notice that the e- is on the right.

In a half reaction, only one element is shown and the charges must be conserved. To write a half reaction, first assign all the oxidation numbers to all the elements. Second, cross out any elements that are spectators (they do not change oxidation number). Then, write a half reaction showing a change in oxidation state and label which element is being oxidized and which is being reduced. Last, add in the number of electrons needed to conserve the charge.

Directions: For the following examples, determine if they represent oxidation or reduction.

1. e- + Cr+3 🡪 Cr+2
2. 2e- + Mn+7 🡪 Mn+5
3. Mg+2 + 2e- 🡪 Mg
4. K 🡪 e- + K+1
5. e- + Cu+2 🡪 Cu+1
6. Al🡪 Al+3 + 3e-
7. Li+ + e- 🡪 Li
8. S-2 🡪 2e- + S

Directions: For the following examples, add in the e- to balance the charge.

1. Cr+5 🡪 Cr+2
2. Mn+4 🡪 Mn+7
3. Ca+2 🡪 Ca
4. Rb 🡪 Rb+1
5. Cu+1 🡪 Cu+2
6. B+3 🡪 B
7. N-3 🡪 N+4
8. Te-2 🡪 Te

# **Spontaneous Reactions and Activity Series**

# Use Table J to help you with the following questions.

1. Write the oxidation and reduction half reaction for:

Ca + Cu2+ 🡪 Ca2+  + Cu

1. According to Table J, the element higher on the list will oxidize. Which element is oxidizing here and does that mean that this reaction is spontaneous?
2. Write the oxidation and reduction half reaction for

Mg + Ca2+ 🡪 Mg2+ + Ca

1. Which element is oxidizing in number 3 and is the reaction spontaneous?
2. Which of the following ions is most easily oxidized?
   1. F- b. Cl- c. Br- d. I-
3. Which element is more easily reduced?
   1. Cu b. Mg c. Al d. Zn
4. Which element will reduce Mg2+ to Mg?
   1. Fe b. Ba c. Pb d. Ag
5. Which ion will oxidize Fe?
   1. Zn2+ b. Ca2+ c. Mg2+ d. Cu2+
6. Which metal will react spontaneously with Ag+ but not Zn2+?
   1. Cu b. Au c. Al d. Mg
7. Which reaction will take place spontaneously?
   1. Mg + Ca2+ 🡪 Mg2+ + Ca
   2. Ba + 2Na+ 🡪 Ba2+ + 2Na
   3. Cl2 + 2F-🡪 2Cl- +F2
   4. I2 + 2Br- 🡪 2I- + Br2

**RedOx Reaction Regents Review**

1. In an oxidation-reduction reaction, reduction is defined as the  
 (1) loss of mass     (3) loss of electrons  
 (2) gain of mass    (4) gain of electrons

2. When a neutral atom undergoes oxidation,

the atom’s oxidation state

1. decreases as it gains electrons
2. decreases as it loses electrons
3. increases as it gains electrons
4. increases as it loses electrons

3. In a redox reaction, there is a conservation of

1. mass, only
2. both mass and charge
3. neither mass nor charge

4. In any redox reaction, the substance that undergoes reduction will

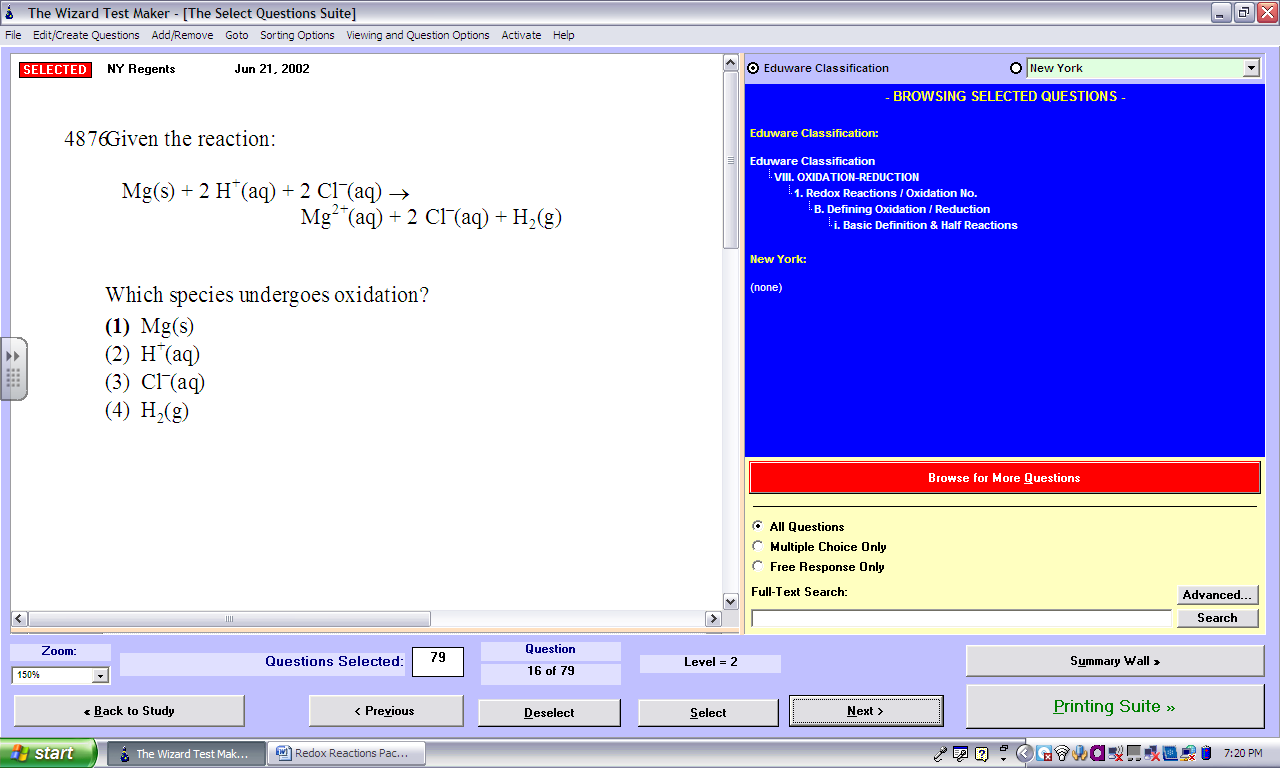
(1) lose e- & have a decrease in oxidation number

(2) lose e- & have an increase in oxidation number

(3) gain e- & have a decrease in oxidation number

(4) gain e- & have an increase in oxidation number

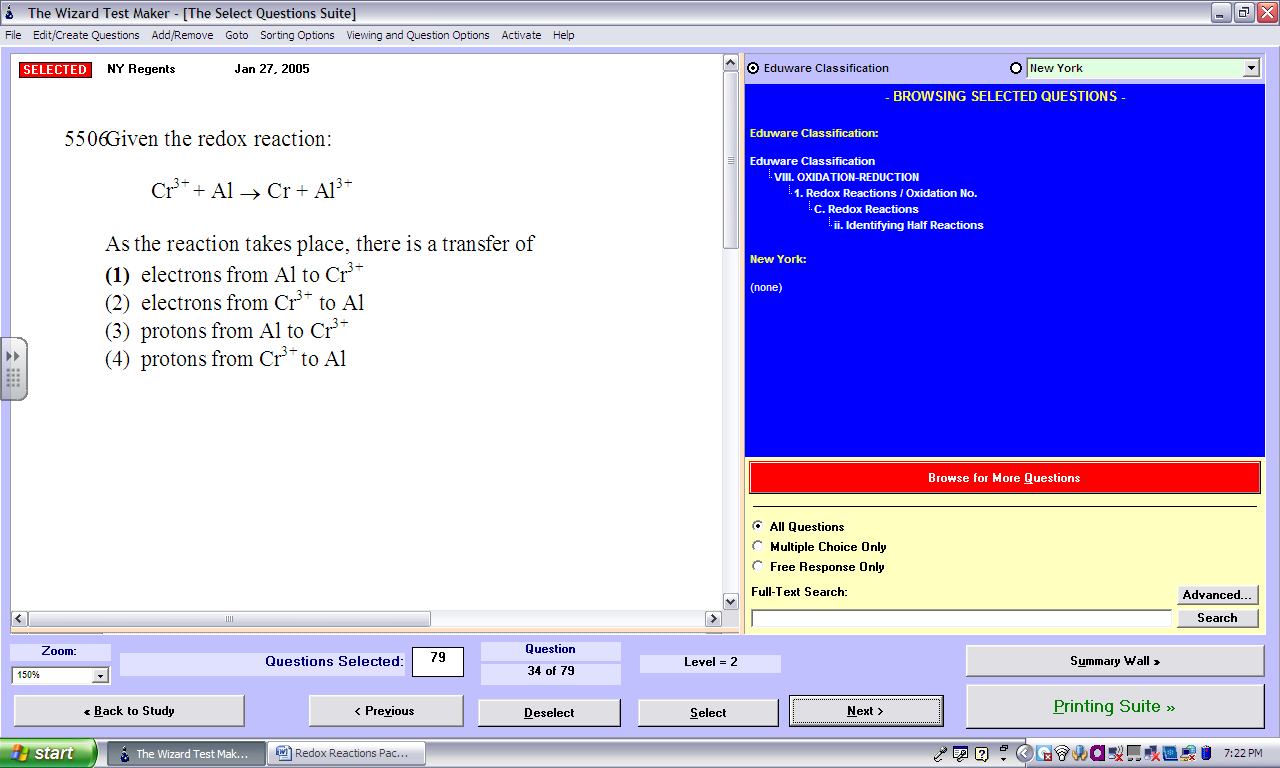
5. Given the reaction:



Which species undergoes oxidation?

1. Mg(s) (3) Cl–(aq)
2. H+(aq) (4) H2(g)

6. Given the redox reaction:



As the reaction takes place, there is a transfer of

1. electrons from Al to Cr3+
2. electrons from Cr3+ to Al

7. Given the redox reaction:

2 Fe3+ + 3 Zn 🡪 2 Fe + 3 Zn2+

As the reaction takes place, there is a transfer of electrons

1. from Fe3+ to Zn
2. from Zn to Fe3+
3. from Zn2+ to Fe

(4) from Fe to Zn2+

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

9. Circle the electrons in the half-reactions below and identify as oxidation or reduction.

(g) Cu2+ + 2 e– → Cu

(h) Fe → Fe2+ + 2 e–

(i) Mn7+ + 3 e– → Mn4+

(d) Cl2 + 2 e– → 2 Cl–

(e) Na+ + e– → Na

(f) S2– → S + 2e–

1. Br2 + 2 e–  → 2 Br–

(b) Na → Na+ + e–

(c) Ca2+ + 2e– → Ca

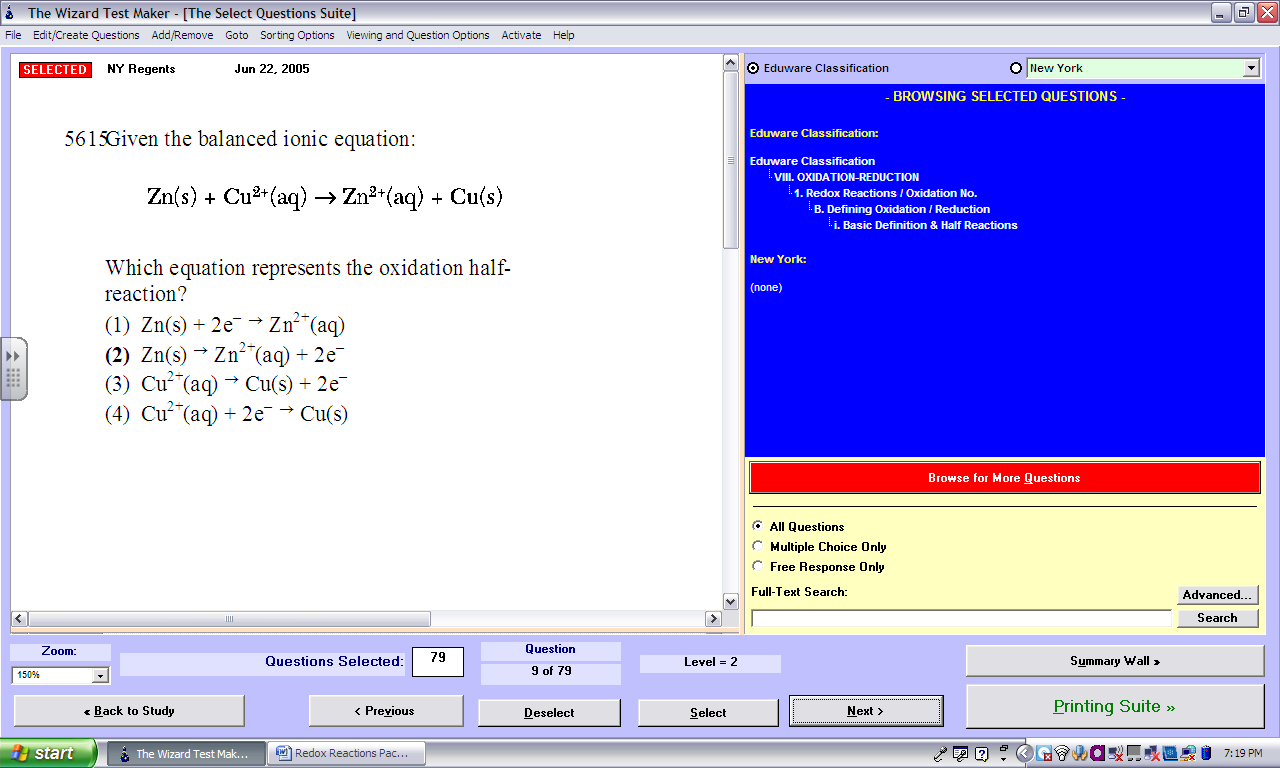
10. Complete the half-reactions below by ADDING in electrons to the correct side in order to equalize charge (show conservation of charge).

1. Fe2+ 🡪 Fe3+
2. K 🡪 K+
3. Sn4+ 🡪 Sn2+
4. Cr6+ 🡪 Cr3+
5. Mn3+ 🡪 Mn4+
6. Cr2+ 🡪 Cr3+
7. Cl7+ 🡪 Cl1+

11. Which half-reaction correctly represents oxidation?

1. Fe(s) → Fe2+(aq) + 2e–
2. Fe2+(aq) → Fe(s) + 2e–
3. Fe(s) + 2e– → Fe2+(aq)
4. Fe2+(aq) + 2e– → Fe(s)

12. Given the balanced ionic equation:



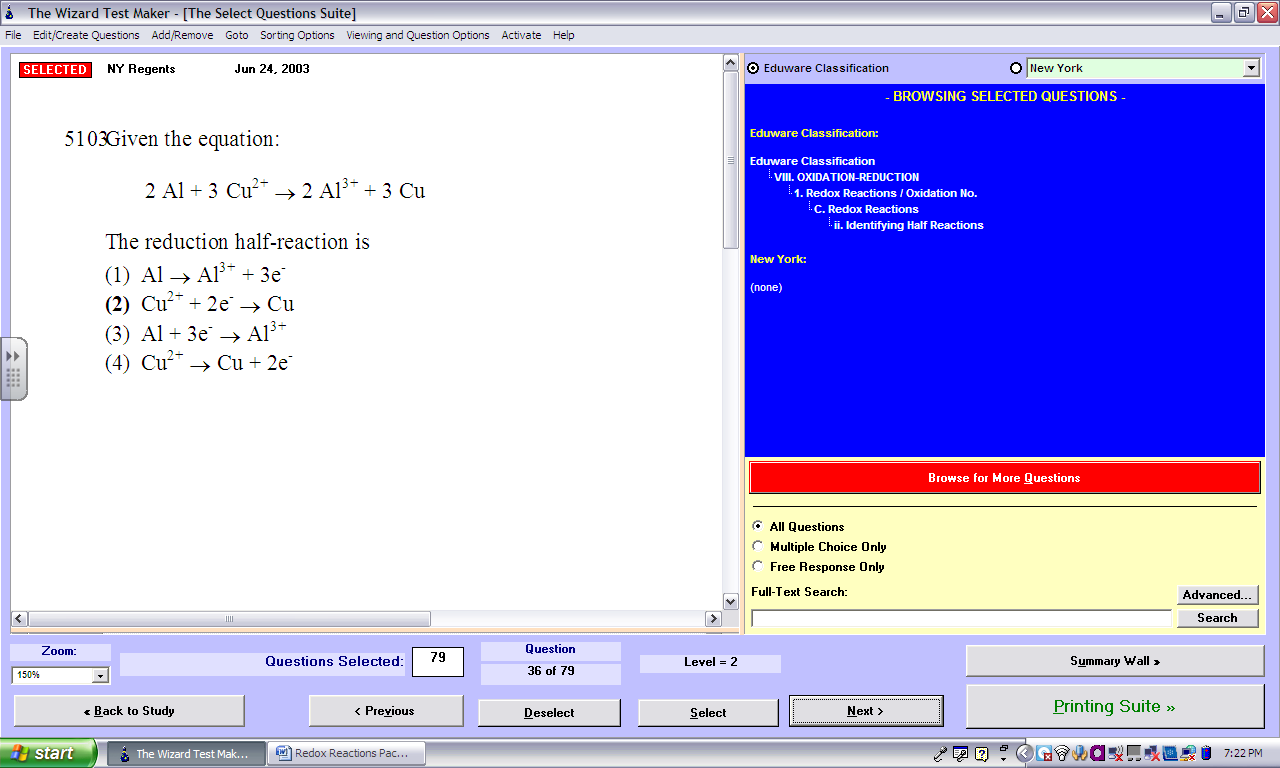
Which equation represents the oxidation

half-reaction?

1. Zn(s) + 2e– → Zn2+(aq)
2. Zn(s) → Zn2+(aq) + 2e–
3. Cu2+(aq) → Cu(s) + 2e–
4. Cu2+(aq) + 2e– → Cu(s)

13. Which half-reaction equation represents the reduction of a potassium ion?  
(1) K+ + e– 🡪 K    (3) K+ 🡪 K + e–  
(2) K + e–🡪 K+    (4) K 🡪 K+ + e–

14. Given the equation:



The reduction half-reaction is

1. Al → Al3+ + 3e–
2. Cu2+ + 2e– → Cu
3. Al + 3e– → Al3+
4. Cu2+ → Cu + 2e–

15. Base your answers to the questions below on the following redox reaction, which occurs in a battery.

\_\_\_\_ Zn + \_\_\_\_ Cr3+ 🡪 \_\_\_\_ Zn2+ + \_­­\_\_\_ Cr

1. Write the half-reaction for the reduction that occurs.
2. Write the half-reaction for the oxidation that occurs.
3. Which species loses electrons and which species gains electrons?

16. The outer structure of the Statue of Liberty is made of copper metal. The framework is made of iron. Over time, a thin green layer (patina) forms on the copper surface.

1. When copper oxidized to form this patina layer, the copper atoms became copper(II) ions (Cu2+). Write a balanced half-reaction for this oxidation of copper.

CCu+ 2e–

1. Where the iron framework came in contact with the copper surface, a reaction occurred in which iron was oxidized. Using information from Reference Table *J*, explain why the iron was oxidized. more reactive than Ni.

**FORMULAS OF IONIC COMPOUNDS**

Ionic Compounds require two types of ions: **cations** which are positive and **anions** which are negative. All **metals** (on the left side of the periodic table) form cations and **nonmetals** (on the left side of the periodic table) form anions primarily. In order to determine the formula of the compound they create you must make sure their ions sum to zero. For example, table salt is sodium chloride. Using the periodic table’s first set of ions, sodium forms +1 ions and chlorine forms -1 ions. Therefore their ions cancel out and the formula is NaCl. It is not always that easy. Calcium chloride is the salt we put on roads to melt ice. Calcium forms +2 ions and Chloride forms -1 ions. We need two chloride ions to balance the charges. The formula is CaCl­2. **Notice the metal, or positive cation is always written first**! Try the following examples:

1. Cesium fluoride: \_\_\_\_\_\_\_\_\_\_\_\_\_\_ 4. Barium sulfide: \_\_\_\_\_\_\_\_\_\_\_\_\_\_

2. Potassium oxide: \_\_\_\_\_\_\_\_\_\_\_\_\_\_ 5. Aluminum chloride: \_\_\_\_\_\_\_\_\_\_\_\_\_\_

3. Magnesium iodide: \_\_\_\_\_\_\_\_\_\_\_\_\_\_ 6. Calcium phosphide: \_\_\_\_\_\_\_\_\_\_\_\_\_\_

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
|  | Chloride | Sulfide | Fluoride | Phosphide |
| Lithium |  |  |  |  |
| Aluminum |  |  |  |  |
| Magnesium |  |  |  |  |
| Potassium |  |  |  |  |

Now we know how to write formulas from their names but we also need to know how to write names from formulas. The rule is: write the whole name of the first element and the second element drop the ending and replace with “ide.” For example: H2S is hydrogen sulfide. In this case, the amount of each element doesn’t affect the name of the compound. Use **table S** to help you find names. Try to name the following examples:

1. NaF \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ 6. NaH \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

2. MgCl2 \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ 7. K3P \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

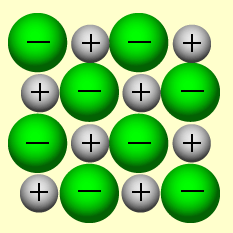
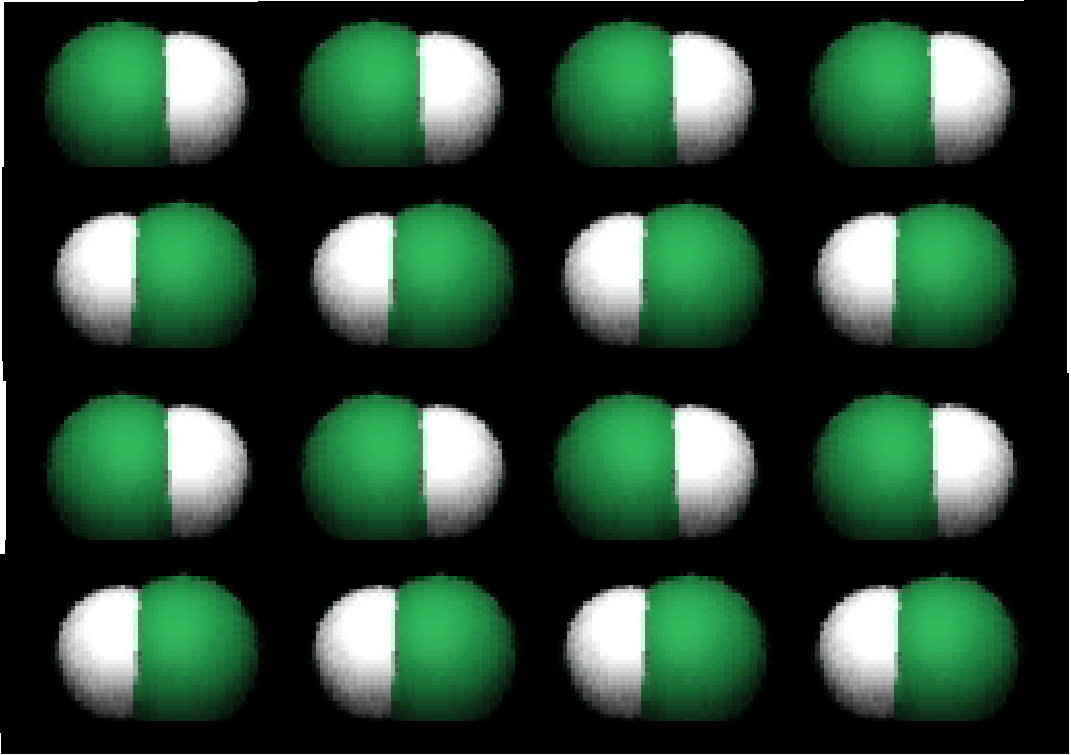
3. Al2O3 \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ 8. MgO \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

4. MgI2 \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ 9. Li2Te \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

5. H2O \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ 10. AlCl3 \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Unit 5 – Review

1. Recall your representations of the atoms in the Sticky Tape activity. Below is a pair of tapes before they have been pulled apart. Explain why they would **not** exert a force (either attractive or repulsive) on one another.
2. Below are groups of the inner cores of the atoms of the tapes after they have been pulled apart. Sketch in the mobile negative charges to show how the top tape becomes (+) and the bottom becomes (-).
3. What evidence allowed us to conclude that the top tape was (+)?
4. Below is a group of the inner cores of a piece of metal foil. Sketch in where you would expect to find the mobile negative charges if a top (+) tape were brought to the left of the foil. Explain your diagram.
5. Describe how JJ Thomson concluded that the mobile charged particle in the atom had a (–) charge.
6. A solution of salt conducts electricity; a solution of sugar does not. Explain.
7. Below left is a 2-D array that represents an ionic lattice. At right is a 2-D array that represents a molecular solid. In what ways are they similar? In what ways are they different?



1. What evidence helped us to conclude that chloride ions have a (–) charge?
2. How do you decide how many ions of each type combine to form an ionic compound?
3. Why do ionic solids have higher melting and boiling points than do most molecular solids?

**RedOX Review Part 2**

Oxidation Number Rules:

1. Elements not in a compound (including diatomic elements e.g. H2) have an oxidation number of 0 because they have not lost or gained electrons yet so they are neutral.
2. Assign the oxidation number to the last element in the compound first (this is negative). Then assign an oxidation number to the first element (this is positive). Remember that the sum of the oxidation numbers of all the elements in a compound equals 0 because (like elements) compounds are neutral.

**Directions Questions 1-8:** Providethe oxidation number of the indicated atoms.

1. Na

1. Na in NaCl

1. H in H2O

1. Ba in BaBr2
2. S in Al2S3

1. N in Na3N

1. Mn in MnO2
2. Br2

1. Explain why sodium is not charged when it is a pure metal but becomes charges when it is in a compound like table salt.
2. Explain the difference between pure silver, in clean jewelry or silverware utensils, versus tarnished silver when jewelry and utensils are reacted with oxygen in the air in terms of charges and electrons.

**Directions Questions 11- 14:** Assign oxidation states to each species below. Identify which species is being oxidized and which is being reduced. In the space below each reaction write the oxidation and reduction half reactions and add in electrons to balance the charge.

oxidation

***Example:*** Li0 + Cu+2 →Li1+ + Cu0

reduction

Oxidization: Li0 →Li1+ + 1e- Reduction: Cu+2 +2e- → Cu0

1. Cr+ + Sn4+ → Cr3+ + Sn2+

Oxidization: Reduction:

1. CaCl2(aq) + Li(s) → LiCl(aq) + Ca(s)

Oxidized: Reduced:

1. As(s) + Cl2(g) → AsCl3(s)

Oxidized: Reduced:

1. Zn(s) + Cu2+(aq) → Zn2+(aq) + Cu(s)

Oxidized: Reduced:

1. Explain how Zn(s) is physically and chemically different than the Zn+2(aq).

**Model 1:** Recall that atoms and compounds are neutral (positive charges = # of negative charges). If element X has 4 negative charges it would also have a positive charge equal in strength to the electrons as shown below.

- -

- -

1. Element X is a metal so it loses electrons (oxidation) to form a cation (positive ion). Draw a model of the atom showing a loss of 2 negative charges (electrons).
2. How strong is the positive charge in the ion (how many electrons is it equal to)?
3. How many negative charges does the ion have?
4. What is the overall charge of the ion?

**Model 2:** Recall that atoms and compounds are neutral (positive charge = # of negative charges). If element Z has 8 negative charges it would also have a positive charge equal in strength to the negative charges as shown below.

- - - -

* - - -

1. Compare Model 1 atom of element X to Model 2 atom of element Z.
2. Which has more electrons?
3. If both atoms of neutral, what can be said about the strength of the positive charge in element Z versus X?
4. How is strength of the positive charge of element Z in Model 2 shown differently than Model 1 of element X?
5. Element Z is a nonmetal so it gains electrons (reduction) to form an anion (negative ion). Draw a model of the atom showing a gain of 2 negative charges (electrons).
6. How strong is the positive charge in the ion (how many electrons is it equal to)?
7. How many negative charges does the ion have?
8. What is the overall charge of the ion?

1. In terms of oxidation state, why is this called reduction when it is gaining electrons?

**Our Model so far…**

Complete the table below using your knowledge of the internal structure of atoms, how they attract and react.

|  |  |
| --- | --- |
| **Diagram:** | **Mathematical:** |
| **Experimental:** | **Narrative:** |